

Chem 103 lecture 1b

Units of concentration Review colligative properties

Units of concentration: M & m

Concentration = ratio of amt solute/amt of solution.

Intrinsic property like density, depends not on total mass

i.e. concentration in 1 drop = concentration in whole pool.

There are several common units of concentration:

A. Molar concentration: (*molarity*)

$$M = \frac{\text{moles solute}}{\text{L of solution}}$$

B. Molal concentration: (*molality*)

$$m = \frac{\text{moles solute}}{\text{Kg of solvent}}$$

Examples of M and m

(1) How many moles of solute are in 40. mL of the .25M solution?

By definition: $\#M = \frac{\# \text{mol solute}}{\text{L solution}}$ Note: $MV = \text{moles}$

$$\# \text{mol solute (M)(L soln)} = (0.25M)(40.\text{mL})(1\text{L}/10^3\text{mL}) = 0.010$$

(2) How many grams of water must be added to 6.0 g glucose to make a solution of 0.40 m ?

$\#m = \# \text{mol glucose}/\text{kg water};$

$$\text{kg water} = \# \text{mol gluc}/\#m = (6.0\text{g})(1\text{mol}/180\text{g}) / (0.40\text{m})$$

$$\text{g water} = (0.0833\text{kg})(10^3\text{g}/1\text{kg}) = 83\text{ g}$$

Other units : %(m/m), %(m/v)

C. Mass percent: ($\%(m/m)$) Sometimes, more conveniently written as:

$$\%(m/m) = \frac{\text{g solute}}{\text{g solvent}} \times 100\% \quad \%(m/m) = \frac{\text{g solute}}{100\text{g solvent}}$$

D. Mass/Vol percent: ($\%(m/v)$) Sometimes, more conveniently written as:

$$\%(m/v) = \frac{\text{g solute}}{\text{mLs of solvent}} \times 100\% \quad \%(m/v) = \frac{\text{g solute}}{100\text{ mLs solvent}}$$

Samples problems: %(m/v)

(1) If you dissolve 20.0 g sucrose (342g/mole) in 150.mL solution, %(m/v)= ?

$$\%(\text{m/v}) \text{ sucrose} = \frac{20.\text{g}}{150.\text{mL}} \times 100\% = 13\%$$

(2) What is the molarity of a 5.20%(m/v) solution of NaCl (FW=58.5g/mol)?

$$\#M = \frac{5.20\text{g}}{100\text{mL}} \frac{1\text{mol}}{58.5\text{g}} \frac{1000\text{mL}}{\text{L}} = 0.889\text{M}$$

Other units: ppm

E. parts per million : *ppm*

Convenient form for dilute aqueous solutions:

$$1 \text{ ppm} = \frac{1\text{gram solute}}{10^6\text{grams solution}} \times \frac{x10^{-3}}{x10^{-3}} = \frac{1x10^{-3}\text{g solute}}{10^3\text{g sol'n}}$$

For dilute aqueous solutions, density is $\approx 1\text{g/mL}$ so, in that case:

$$1\text{ppm} = \frac{1 \text{ mg solute}}{1\text{L solution}} \quad 1 \text{ ppm} = \frac{1x10^{-3}\text{g solute}}{10^3\text{g sol'n}} \times \frac{1\text{g}}{1\text{mL}} = \frac{1 \text{ mg solute}}{1\text{L solution}}$$

But for gases:

$$1\text{ppm} = \frac{1 \text{ mole gas solute}}{10^6\text{mole gas mixture}}$$

Ppm for gases

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F. parts per billion: *ppb*

For dilute aqueous solutions:

$$1\text{ppm} = \frac{1 \mu\text{g solute}}{1\text{L solution}}$$

Sample problem: ppm

How many ppm Na^+ (23g/mol) is present in a 5.0×10^{-4} M NaCl (58.5g/mol) solution?

Solution: remember that #ppm = # mg/L for this solution.

$$\#\text{ppm Na}^+ = \frac{5.0 \times 10^{-4} \text{ mol NaCl}}{\text{L}} \frac{1 \text{ mol Na}^+}{1 \text{ mol NaCl}} \frac{23 \text{ g Na}}{1 \text{ mol Na}} \frac{10^3 \text{ mg}}{\text{g}}$$

$$\#\text{ppm Na}^+ = 11.5 \text{ mg/L} = 11.5 \text{ ppm}$$

Other units: mole fraction, X

Mole fraction, X_i of a solute “i” is defined as:

$$X_i = \text{moles } i / (\text{total moles of solution})$$

For example, if you mix 2.0 moles of naphthalene to 3.0 moles of hexane, you will get a solution whose total moles = $2.0 + 3.0 = 5.0$ moles.

$$X_{\text{naphthalene}} = 2.0 \text{ moles} / 5.00 \text{ moles} = 0.400$$

Colligative properties

What are colligative properties?

Physical properties due to *collective* effect of solute molecules

In other words, due to *number* of solute molecules or ions

Does **NOT** depend on the identity of the solute.

Applies to electrolyte and nonelectrolyte solutions

4 common colligative properties

4 common colligative properties are:

- 1) Vapor pressure **lowering**
- 2) Boiling point **elevation**
- 3) Freezing point **depression**
- 4) **Osmotic pressure**

Vapor pressure lowering

$$\text{Raoult's Law: } P_1 = X_1 P_1^\circ$$

where P_1 = vapor pressure of solvent.

1 = solvent.

$$X_1 = \text{mol fraction of solvent} = n_1 / (n_1 + n_2 + \dots)$$

Example: vapor pressure of salt water is less than that of distilled water at equilibrium.

Consequence of Raoult's Law

What would happen if we place 2 beakers (one half-filled with pure water and the other half-filled with sugar water) in an enclosed chamber as follows:

